CHAPTER 4

The Structure of the Atom

What You'll Learn

- You will identify the experiments that led to the development of the nuclear model of atomic structure.
- You will describe the structure of the atom and differentiate among the subatomic particles that comprise it.
- You will explain the relationship between nuclear stability and radioactivity.

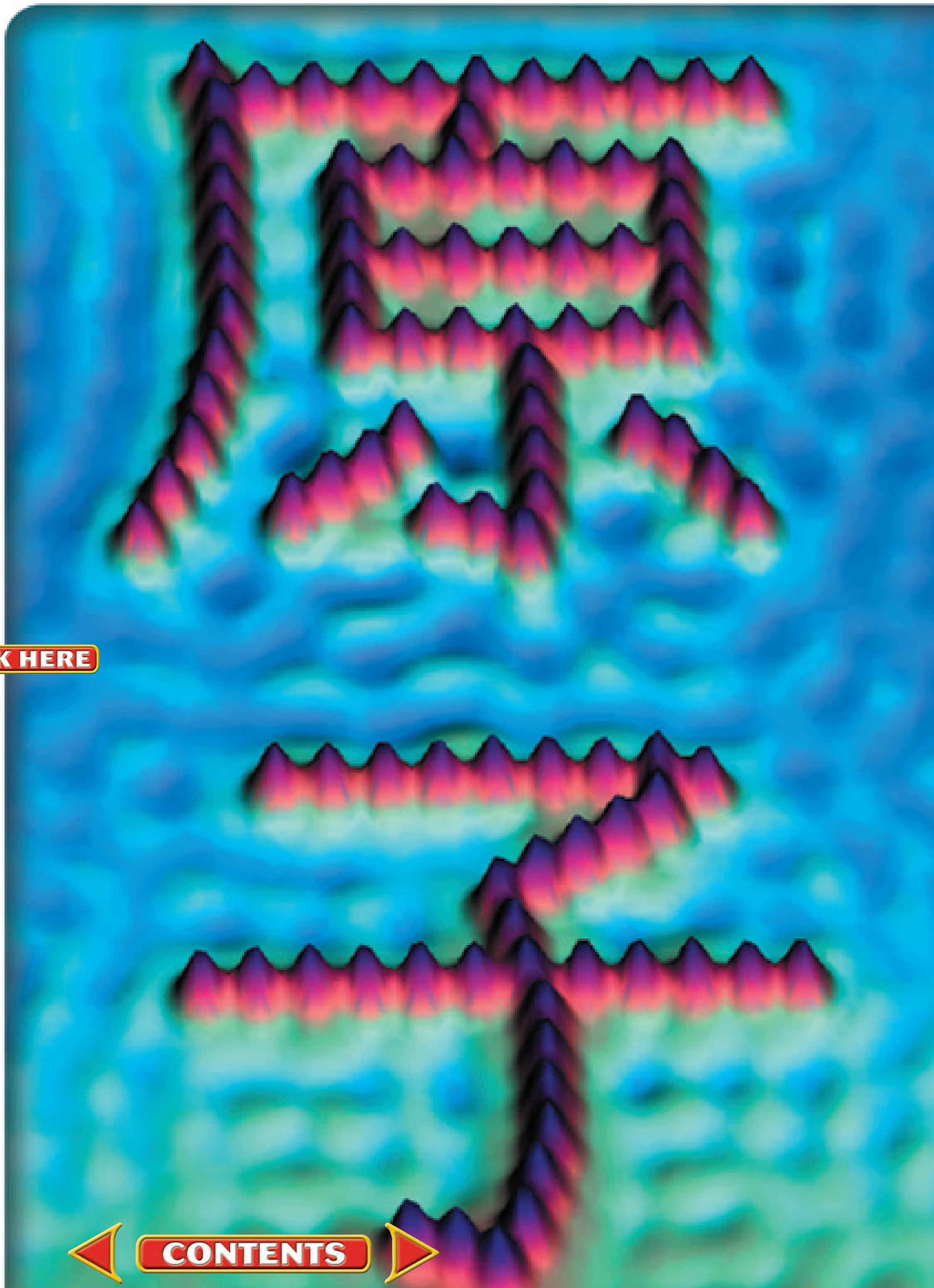
Why It's Important

The world you know is made of matter, and all matter is composed of atoms. Understanding the structure of the atom is fundamental to understanding why matter behaves the way it does.

CLICK HERE

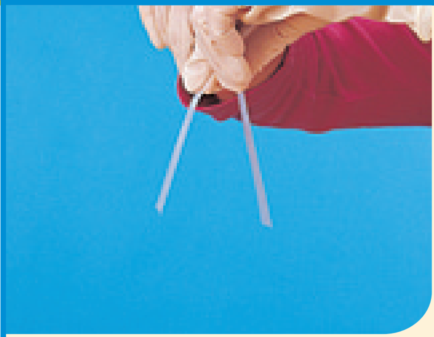
CHEMISTRY Online

Visit the Chemistry Web site at **science.glencoe.com** to find links about atomic structure.



Not only can individual atoms be seen, but scientists now have the ability to arrange them into patterns and simple devices. The atoms shown here have been arranged to form the Japanese kanji characters for atom.

DISCOVERY LAB



Observing Electrical Charge

Electrical charge plays an important role in atomic structure and throughout chemistry. How can you observe the behavior of electrical charge using common objects?

Materials

metric ruler	10-cm long piece of clear plastic tape (4)
plastic comb	
hole punch	
paper	

Procedure

1. Cut out small round pieces of paper using the hole punch and spread them out on a table. Run a plastic comb through your hair. Bring the comb close to the pieces of paper. Record your observations.
2. Fold a 1-cm long portion of each piece of tape back on itself to form a handle. Stick two pieces of tape firmly to your desktop. Quickly pull both pieces of tape off of the desktop and bring them close together so that their non-sticky sides face each other. Record your observations.
3. Firmly stick one of the remaining pieces of tape to your desktop. Firmly stick the last piece of tape on top of the first. Quickly pull the pieces of tape as one from the desktop and then pull them apart. Bring the two tape pieces close together so that their non-sticky sides face each other. Record your observations.

Analysis

Use your knowledge of electrical charge to explain your observations. Which charges are similar? Which are different? How do you know?

Section 4.1 Early Theories of Matter

Objectives

- **Compare** and **contrast** the atomic models of Democritus and Dalton.
- **Define** an atom.

Vocabulary

Dalton's atomic theory
atom

Perhaps you have never seen a photo of individual atoms as shown on the previous page, but chances are you've heard of atoms ever since you were in elementary school. From atom smashers and atomic power to the reality of the atomic bomb, you are already familiar with many modern atom-based processes. Surprisingly, the idea that matter is composed of tiny particles (which we now call atoms) did not even exist a few thousand years ago. In fact, for more than a thousand years, great thinkers of their day argued against the idea that atoms existed. They were wrong. As you will see, the development of the concept of the atom and our understanding of atomic structure are fascinating stories involving scores of great thinkers and scientists.

The Philosophers

Science as we know it today did not exist several thousand years ago. No one knew what a controlled experiment was, and there were few tools for scientific exploration. In this setting, the power of mind and intellectual thought were considered the primary avenues to the truth. Curiosity sparked the interest of scholarly thinkers known as philosophers who considered the many mysteries of life. As they speculated about the nature of matter, many of the philosophers formulated explanations based on their own life experiences.

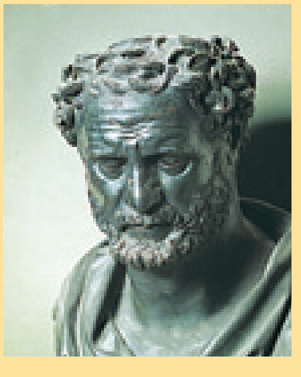
Figure 4-1

Many Greek philosophers thought matter was formed of air, earth, fire, and water. They also associated properties with each of the four basic components of matter. The pairings of opposite properties, such as hot and cold, and wet and dry, mirrored the symmetry and balance the philosophers observed in nature. These early nonscientific and incorrect beliefs were not completely dispelled until the 1800s.

It wasn't surprising then, that many of them concluded that matter was composed of things such as earth, water, air, and fire. See **Figure 4-1.** It was also commonly accepted that matter could be endlessly divided into smaller and smaller pieces. While these early ideas were creative, there was no method for testing their validity.

The Greek philosopher Democritus (460–370 B.C.) was the first person to propose the idea that matter was not infinitely divisible. He believed matter was made up of tiny individual particles called *atomos,* from which the English word atom is derived. Democritus believed that atoms could not be created, destroyed, or further divided. Democritus and a summary of his ideas are shown in **Figure 4-2**.

While a fair amount of Democritus's ideas do not agree with modern atomic theory, his belief in the existence of atoms was amazingly ahead of his time. Despite this, his ideas did not turn out to be a major step toward our current understanding of matter. Over time, Democritus's ideas were met with criticism from other philosophers. "What holds the atoms together?" they asked. Democritus could not answer the question. Other criticisms came from Aristotle (384–322 B.C.), one of the most influential Greek philosophers. Aristotle is shown in **Figure 4-3**. He rejected the atomic "theory" entirely



Democritus's Ideas

- Matter is composed of empty space through which atoms move.
- Atoms are solid, homogeneous, indestructible, and indivisible.
- Different kinds of atoms have different sizes and shapes.
- The differing properties of matter are due to the size, shape, and movement of atoms.
- Apparent changes in matter result from changes in the groupings of atoms and not from changes in the atoms themselves.

Figure 4-2

The Greek philosopher Democritus (460–370 B.C.) proposed the concept of the atom more than two thousand years ago.

Aristotle

- One of the most influential philosophers.
- Wrote extensively on many subjects, including politics, ethics, nature, physics, and astronomy.
- Most of his writings have been lost through the ages.

Figure 4-3

The Greek philosopher Aristotle (384–322 B.C.) was influential in the rejection of the concept of the atom.

because it did not agree with his own ideas on nature. One of Aristotle's major criticisms concerned the idea that atoms moved through empty space. He did not believe that the "nothingness" of empty space could exist. Unable to answer the challenges to his ideas, Democritus's atomic theory was eventually rejected.

In fairness to Democritus, it was impossible for him or anyone else of his time to determine what held the atoms together. More than two thousand years would pass before the answer was known. However, it is important to realize that Democritus's ideas were just that—ideas and not science. Without the benefit of being able to conduct controlled experiments, Democritus could not test to see if his ideas were valid.

Unfortunately for the advancement of science, Aristotle was able to gain wide acceptance for his ideas on nature—ideas that denied the existence of atoms. Incredibly, the influence of Aristotle was so great and the development of science so primitive that his denial of the existence of atoms went largely unchallenged for two thousand years!

CHEMISTRY CD-ROM

Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

John Dalton

Although the concept of the atom was revived in the 18th century, it took the passing of another hundred years before significant progress was made. The work done in the 19th century by John Dalton (1766–1844), a schoolteacher in England, marks the beginning of the development of modern atomic theory. Dalton revived and revised Democritus's ideas based upon the results of scientific research he conducted. The main points of **Dalton's atomic theory** are shown in **Figure 4-4**.

Dalton's Atomic Theory

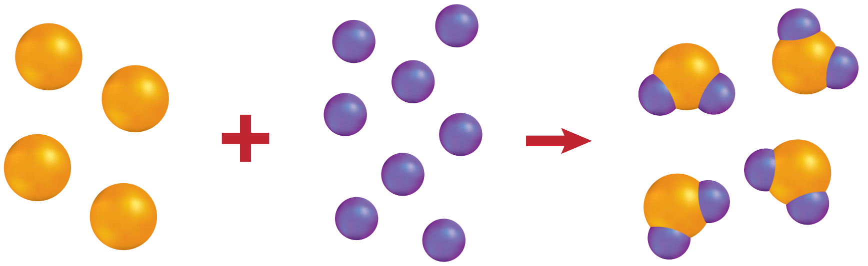
- All matter is composed of extremely small particles called atoms.
- All atoms of a given element are identical, having the same size, mass, and chemical properties. Atoms of a specific element are different from those of any other element.
- Atoms cannot be created, divided into smaller particles, or destroyed.
- Different atoms combine in simple whole-number ratios to form compounds.
- In a chemical reaction, atoms are separated, combined, or rearranged.

Figure 4-4

John Dalton's (1766–1844) atomic theory was a breakthrough in our understanding of matter.

Figure 4-5

Dalton's atomic theory explains the conservation of mass when a compound forms from its component elements. Atoms of elements A and B combine in a simple whole-number ratio, in this case two B atoms for each A atom, to form a compound. Because the atoms are only rearranged in the chemical reactions, their masses are conserved.



Atoms of element A	Atoms of element B	Compound composed of elements A and B
Total mass = mass A	Total mass = mass B	Total mass = (mass A + mass B)

The advancements in science since Democritus's day served Dalton well, as he was able to perform experiments that allowed him to refine and verify his theories. Dalton studied numerous chemical reactions, making careful observations and measurements along the way. He was able to accurately determine the mass ratios of the elements involved in the reactions. Based on this research, he proposed his atomic theory in 1803. In many ways Democritus's and Dalton's theories are similar. What similarities and differences can you find between the two theories?

Recall from Chapter 3 that the law of conservation of mass states that mass is conserved in any process, such as a chemical reaction. Dalton's atomic theory easily explains the conservation of mass in chemical reactions as being the result of the separation, combination, or rearrangement of atoms—atoms that are not created, destroyed, or divided in the process. The formation of a compound from the combining of elements and the conservation of mass during the process are shown in **Figure 4-5**. Dalton's convincing experimental evidence and clear explanation of the composition of compounds and conservation of mass led to the general acceptance of his atomic theory.

Was Dalton's atomic theory a huge step toward our current atomic model of matter? Yes. Was all of Dalton's theory accurate? No. As is often the case in science, Dalton's theory had to be revised as additional information was learned that could not be explained by the theory. As you will soon learn, Dalton was wrong about atoms being indivisible (they are divisible into several subatomic particles) and about all atoms of a given element having identical properties (atoms of an element may have slightly different masses).

History CONNECTION

Alchemy, a popular pursuit during the Middle Ages, was the search for a way to transform common metals into gold. "Scientists" who studied alchemy did not succeed in producing gold, but their experiments revealed the properties of various metals and helped advance our understanding of matter. During the seventeenth century some alchemists began focusing on identifying new compounds and reactions—these alchemists were probably the first true chemists. By the twentieth century, scientists could transform atoms of some elements into atoms of other elements by bombarding them with neutrons. Even this process, however, has not succeeded in producing gold.

Defining the Atom

Many experiments since Dalton's time have proven that atoms do actually exist. So what exactly then is the definition of an atom? To answer this question, consider a gold ring. Suppose you decide to grind the ring down into a pile of gold dust. Each fragment of gold dust still retains all of the properties of gold. If it were possible—which it is not without special equipment—you could continue to divide the gold dust particles into still smaller particles. Eventually you would encounter a particle that could not be divided any further and still retain the properties of gold. This smallest particle of an element that retains the properties of the element is called an **atom**.

Just how small is a typical atom? To get some idea of its size, consider the population of the world. In the year 2000, the world population was approximately 6 000 000 000 (six billion) people. By comparison, a typical solid copper penny contains almost five billion times as many atoms of copper!

World population	6 000 000 000
Atoms in a penny	29 000 000 000 000 000 000 000

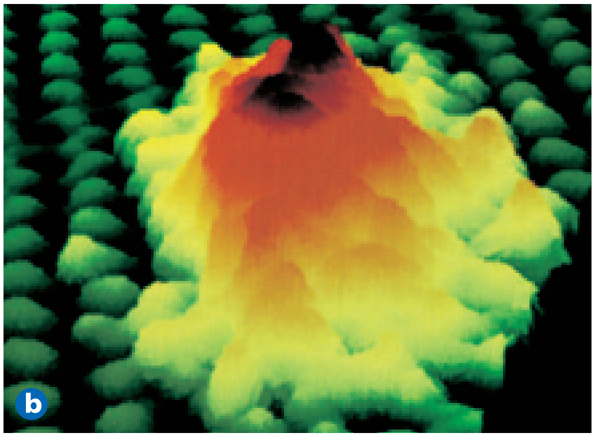


Figure 4-6

a This colorized scanning electron micrograph shows a microgear mechanism built by the precise arrangement of individual atoms.
b A mound of gold atoms (yellow, red, and brown) is easily discerned from the graphite substrate (green) it rests on.

The diameter of a single copper atom is 1.28×10^{-10}m. Placing six billion copper atoms (equal in number to the world's population) side by side would result in a line of copper atoms less than one meter long.

You might think that because atoms are so small there would be no way to actually see them. However, an instrument called the scanning tunneling microscope allows individual atoms to be seen. Do the **problem-solving LAB** on page 96 to analyze scanning tunneling microscope images and gain a better understanding of atomic size. As **Figures 4-6a** and **4-6b** illustrate, not only can individual atoms be seen, scientists are now able to move individual atoms around to form shapes, patterns, and even simple machines. This capability has led to the exciting new field of nanotechnology. The promise of nanotechnology is molecular manufacturing—the atom-by-atom building of machines the size of molecules. As you'll learn in later chapters, a molecule is a group of atoms that are bonded together and act as a unit. While this technology is not yet feasible for the production of consumer products, progress toward that goal has been made. To learn more about nanotechnology, read the **Chemistry and Society** at the end of this chapter.

The acceptance of atomic theory was only the beginning of our understanding of matter. Once scientists were fairly convinced of the existence of atoms, the next set of questions to be answered emerged. What is an atom like? How are atoms shaped? Is the composition of an atom uniform throughout, or is it composed of still smaller particles? While many scientists researched the atom in the 1800s, it was not until almost 1900 that answers to some of these questions were found. The next section explores the discovery of subatomic particles and the further evolution of atomic theory.

Section 4.1 Assessment

1. Why were Democritus's ideas rejected by other philosophers of his time?
2. Define an atom using your own words.
3. Which statements in Dalton's original atomic theory are now considered to be incorrect? Describe how modern atomic theory differs from these statements.
4. **Thinking Critically** Democritus and Dalton both proposed the concept of atoms. Describe the method each of them used to reach the conclusion that atoms existed. How did Democritus's method hamper the acceptance of his ideas?
5. **Comparing and Contrasting** Compare and contrast the atomic theories proposed by Democritus and John Dalton.

Section 4.2

Subatomic Particles and the Nuclear Atom

Objectives

- **Distinguish** between the subatomic particles in terms of relative charge and mass.
- **Describe** the structure of the nuclear atom, including the locations of the subatomic particles.

Vocabulary

cathode ray
electron
nucleus
proton
neutron

In 1839, American inventor Charles Goodyear accidentally heated a mixture of natural rubber and sulfur. The resulting reaction greatly strengthened the rubber. This new rubber compound revolutionized the rubber industry and was eventually used in the manufacturing of automobile tires. Accidental discoveries such as this have occurred throughout the history of science. Such is the case with the discovery of subatomic particles, the particles that make up atoms.

Discovering the Electron

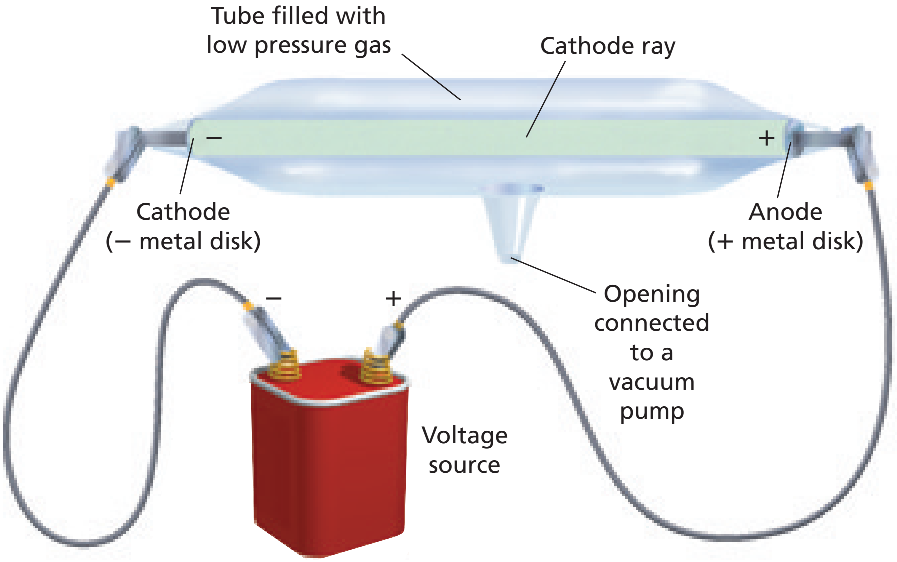
Has your hair ever clung to your comb? Have you ever received a shock from a metal doorknob after walking across a carpeted floor? Observations such as these led scientists in the 1800s to look for some sort of relationship between matter and electric charge. To explore the connection, some scientists wondered how electricity might behave in the absence of matter. With the help of the recently invented vacuum pump, they passed electricity through glass tubes from which most of the air (and most of the matter) had been removed.

A typical tube used by researchers studying the relationship between mass and charge is illustrated in **Figure 4-7.** Note that metal electrodes are located on opposite ends of the tube. The electrode connected to the negative terminal of the battery is called the cathode, and the electrode connected to the positive terminal is called the anode.

One day while working in a darkened laboratory, English physicist Sir William Crookes noticed a flash of light within one of the tubes. The flash was produced by some form of radiation striking a light-producing coating that had been applied to the end of the tube. Further work showed there were rays (radiation) traveling from the cathode to the anode within the tube. Because the ray of radiation originated from the cathode end of the tube, it became known as a **cathode ray.** The accidental discovery of the cathode ray led to the invention of one of the most important technological and social developments of the 20th century—the television. Television and computer monitor images are formed as radiation from the cathode strikes light-producing chemicals that coat the backside of the screen.

Figure 4-7

Examine the parts of a typical cathode ray tube. Note that the electrodes take on the charge of the battery terminal to which they are connected. The cathode ray travels from the cathode to the anode.



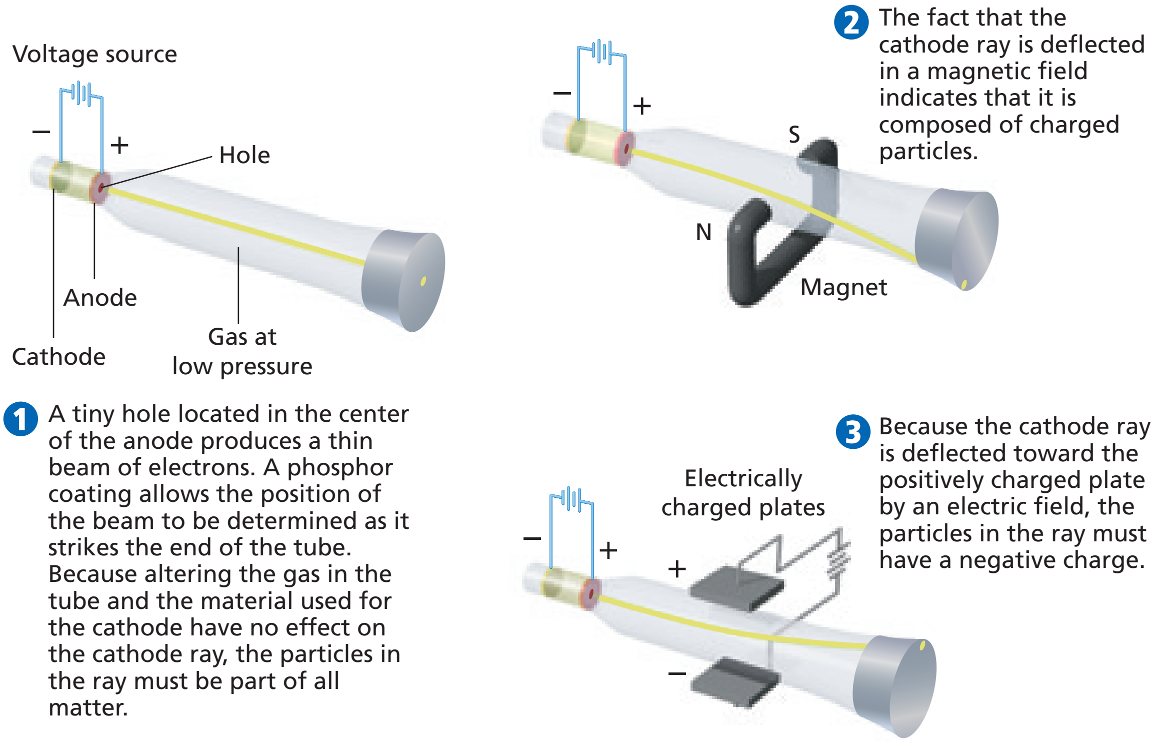


Figure 4-8

Multiple experiments helped determine the properties of cathode rays.

Scientists continued their research using cathode ray tubes, and by the end of the 1800s they were fairly convinced of the following:

- Cathode rays were actually a stream of charged particles.
- The particles carried a negative charge. (The exact value of the negative charge was not known, however.)

Because changing the type of electrode or varying the gas (at very low pressure) in the cathode ray tube did not affect the cathode ray produced, it was concluded that the ray's negative particles were found in all forms of matter. These negatively charged particles that are part of all forms of matter are now called **electrons**. The range of experiments used to determine the properties of the cathode ray are shown in **Figure 4-8**.

In spite of the progress made from all of the cathode ray tube experiments, no one had succeeded in determining the mass of a single cathode ray particle. Unable to measure the particle's mass directly, English physicist J.J. Thomson (1856–1940) began a series of cathode ray tube experiments in the late 1890s to determine the ratio of its charge to its mass. By carefully measuring the effect of both magnetic and electric fields on a cathode ray, Thomson was able to determine the charge-to-mass ratio of the charged particle. He then compared that ratio to other known ratios. Thomson concluded that the mass of the charged particle was much less than that of a hydrogen atom, the lightest known atom. The conclusion was shocking because it meant there were particles smaller than the atom. In other words, Dalton was wrong: Atoms were divisible into smaller subatomic particles. Because Dalton's atomic theory had become so widely accepted, and because Thomson's conclusion was so revolutionary, many fellow scientists found it hard to believe this new discovery. But Thomson was correct. He had identified the first subatomic particle—the electron.

The next significant development came in 1909, when an American physicist named Robert Millikan (1868–1953) determined the charge of an electron.

So good was Millikan's experimental setup and technique that the charge he measured almost one hundred years ago is within 1% of the currently accepted value. This charge has since been equated to a single unit of negative charge; in other words, a single electron carries a charge of 1–. Knowing the electron's charge and using the known charge-to-mass ratio, Millikan calculated the mass of a single electron.

$$\text{Mass of an electron} = 9.1 \times 10^{-28}\ \text{g} = \frac{1}{1840}\ \text{mass of a hydrogen atom}$$

As you can see, the mass of an electron is extremely small.

The existence of the electron and the knowledge of some of its properties raised some interesting new questions about the nature of atoms. It was known that matter is neutral. You know matter is neutral from everyday experience; you do not receive an electrical shock (except under certain conditions) when you touch an object. If electrons are part of all matter and they possess a negative charge, how is it that all matter is neutral? Also, if the mass of an electron is so extremely small, what accounts for the rest of the mass in a typical atom?

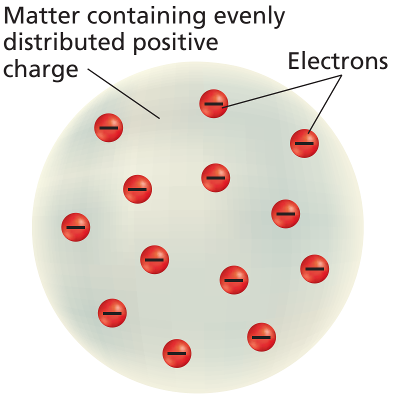


Figure 4-9

J.J. Thomson's plum pudding atomic model proposed that negatively charged electrons were distributed throughout a uniform positive charge.

In an attempt to answer these questions, J.J. Thomson proposed a model of the atom that became known as the plum pudding model. As you can see in **Figure 4-9**, Thomson's model consisted of a spherically shaped atom composed of a uniformly distributed positive charge within which the individual negatively charged electrons resided. A more modern name for this model might be the chocolate-chip cookie dough model, where the chocolate chips are the electrons and the dough is the uniformly distributed positive charge. As you are about to learn, the plum pudding model of the atom did not last for very long.

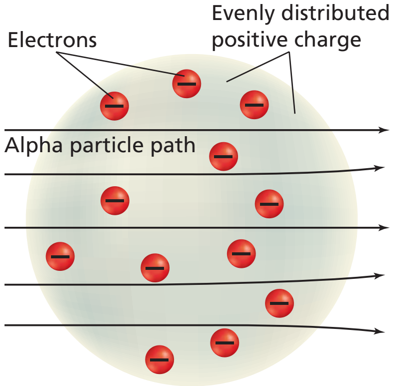
The Nuclear Atom

The story of the atom continues with the role played by Ernest Rutherford (1871–1937). As a youth, Rutherford, who was born in New Zealand, placed second in a scholarship competition to attend the prestigious Cambridge University in England. He received a fortunate break when the winner of the competition decided not to attend. By 1908, Rutherford won the Nobel Prize in chemistry and had many significant discoveries to his credit.

In 1911 Rutherford became interested in studying how positively charged alpha particles (radioactive particles you will learn more about later in this chapter) interacted with solid matter. A small group of scientists that included Rutherford designed and conducted an experiment to see if alpha particles would be deflected as they passed through a thin foil of gold. In the experiment, a narrow beam of alpha particles was aimed at a thin sheet of gold foil. A zinc sulfide coated screen surrounding the gold foil produced a flash of light whenever it was struck by an alpha particle. By noting where the flashes occurred, the scientists could determine if the atoms in the gold foil deflected the alpha particles.

Figure 4-10

Ernest Rutherford expected most of the fast-moving and relatively massive alpha particles to pass straight through the gold atoms. He also expected a few of the alpha particles to be slightly deflected by the electrons in the gold atoms.



Rutherford was aware of Thomson's plum pudding model of the atom and expected only minor deflections of the alpha particles. He thought the paths of the massive (relative to electrons) and fast-moving alpha particles would be only slightly altered by a nearby encounter or collision with an electron. And because the positive charge within the gold atoms was thought to be uniformly distributed, he thought it would not alter the paths of the alpha particles either. **Figure 4-10** shows the results Rutherford anticipated from the experiment. After a few days of testing, Rutherford and his fellow scientists were amazed to discover that a few of the alpha particles were deflected at

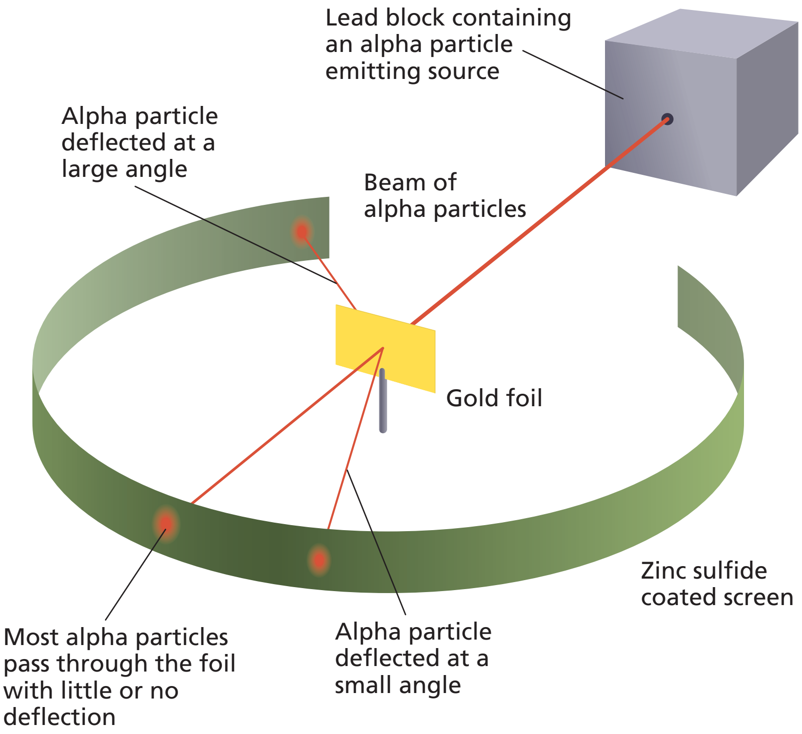


Figure 4-11

As Rutherford expected, most all of the alpha particles passed straight through the gold foil, without deflection. Surprisingly, however, some alpha particles were scattered at small angles, and on a few occasions they were deflected at very large angles.

very large angles. Several particles were even deflected straight back toward the source of the alpha particles. Rutherford likened the results to firing a large artillery shell at a sheet of paper and having the shell come back and hit you! These results, shown in **Figure 4-11**, were truly astounding.

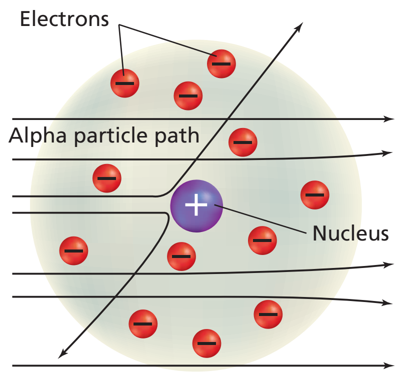
Rutherford concluded that the plum pudding model was incorrect because it could not explain the results of the gold foil experiment. He set out to develop a new atomic model based upon his findings. Considering the properties of the alpha particles and the electrons, and the frequency of the deflections, he calculated that an atom consisted mostly of empty space through which the electrons move. He also concluded that there was a tiny, dense region, which he called the **nucleus**, centrally located within the atom that contained all of an atom's positive charge and virtually all of its mass. Because the nucleus occupies such a small space and contains most of an atom's mass, it is incredibly dense. Just how dense? If a nucleus were the size of the dot in the exclamation point at the end of this sentence, its mass would be approximately as much as that of 70 automobiles!

According to Rutherford's new nuclear atomic model, most of an atom consists of electrons moving rapidly through empty space. The electrons move through the available space surrounding the nucleus and are held within the atom by their attraction to the positively charged nucleus. The volume of space through which the electrons move is huge compared to the volume of the nucleus. A typical atom's diameter, which is defined by the volume of space through which the electrons move, is approximately 10 000 times the diameter of the nucleus. To put this in perspective, if an atom had a diameter of two football fields, the nucleus would be the size of a nickel!

The concentrated positive charge in the nucleus explains the deflection of the alpha particles—the repulsive force produced between the positive nucleus and the positive alpha particles causes the deflections. Alpha particles closely approaching the nucleus were deflected at small angles, while alpha particles directly approaching the nucleus were deflected at very large angles. You can see in **Figure 4-12** how Rutherford's nuclear atomic model explained the

Figure 4-12

Rutherford's nuclear model of the atom explains the results of the gold foil experiment. Most alpha particles pass straight through, being only slightly deflected by electrons, if at all. The strong force of repulsion between the positive nucleus and the positive alpha particles causes the large deflections.



results of the gold foil experiment. The nuclear model also explains the neutral nature of matter: the positive charge of the nucleus balancing the negative charge of the electrons. However, the model still could not account for all of the atom's mass. Another 20 years would pass before this mystery was solved.

Completing the Atom—The Discovery of Protons and Neutrons

By 1920, eight years after his revolutionary gold foil experiment, Rutherford had refined the concept of the nucleus. He concluded that the nucleus contained positively charged particles called protons. A **proton** is a subatomic particle carrying a charge equal to but opposite that of an electron; that is, a proton has a positive charge of 1+.

In 1932, Rutherford's coworker, English physicist James Chadwick (1891–1974), showed that the nucleus also contained another subatomic particle, a neutral particle called the neutron. A **neutron** has a mass nearly equal to that of a proton, but it carries no electrical charge. Thus, three subatomic particles are the fundamental building blocks from which all atoms are

problem-solving LAB

Interpreting STM Images

Measuring The invention of the scanning tunneling microscope (STM) in 1981 gave scientists the ability to visualize individual atoms, and also led to their being able to manipulate the positions of individual atoms. Use the information shown in the STM images to interpret sizes and make measurements.

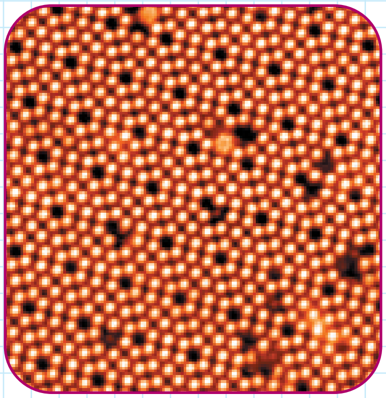


Figure A

Analysis

Figure A is an STM image of silicon atoms that have been bonded together in a hexagonal pattern. The image is of an area 18.1nm wide by 19.0 nm high (1 nm = 1 x 10^{-9} m).
Figure B is an STM image of 48 iron atoms that have been arranged into a circular "corral." The corral has a diameter of 1426 nm. There is a single electron trapped inside the "corral."

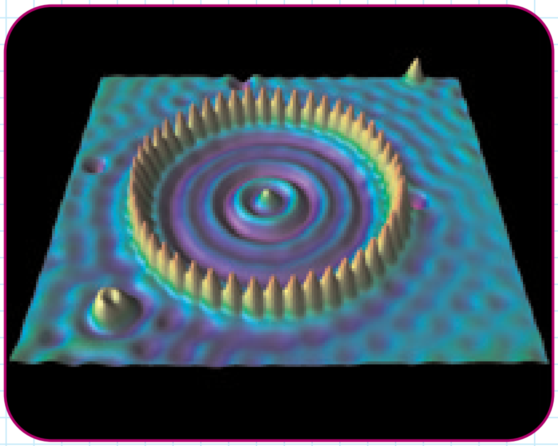


Figure B

Thinking Critically

1. Using a metric ruler and the dimensions of **Figure A** given above, develop a scale for making measurements off of the image. Use your scale to estimate the distance between adjacent silicon nuclei forming a hexagon.
2. What evidence is there that an electron is trapped inside the "corral" of iron atoms in **Figure B**? Estimate the distance between adjacent iron atoms. (Hint: Use the number of atoms and the formula circumference = $\pi \times$ diameter.)

Table 4-1

Properties of Subatomic Particles					
Particle	**Symbol**	**Location**	**Relative electrical charge**	**Relative mass**	**Actual mass (g)**
Electron	e^-	In the space surrounding the nucleus	1−	$\frac{1}{1840}$	9.11×10^{-28}
Proton	p^+	In the nucleus	1+	1	1.673×10^{-24}
Neutron	n^0	In the nucleus	0	1	1.675×10^{-24}

made—the electron, the proton, and the neutron. Together, electrons, protons, and neutrons account for all of the mass of an atom. The properties of electrons, protons, and neutrons are summarized in **Table 4-1**.

You know an atom is an electrically neutral particle composed of electrons, protons, and neutrons. Atoms are spherically shaped, with a tiny, dense nucleus of positive charge surrounded by one or more negatively charged electrons. Most of an atom consists of fast-moving electrons traveling through the empty space surrounding the nucleus. The electrons are held within the atom by their attraction to the positively charged nucleus. The nucleus, which is composed of neutral neutrons (hydrogen's single-proton nucleus is an exception) and positively charged protons, contains all of an atom's positive charge and 99.97% of its mass. Since an atom is electrically neutral, the number of protons in the nucleus equals the number of electrons surrounding the nucleus. The features of a typical atom are shown in **Figure 4-13**. To gain more perspective on the size of typical atoms, do the **CHEMLAB** at the end of this chapter.

Subatomic particle research is still a major interest of modern scientists. In fact, the three subatomic particles you have just learned about have since been found to have their own structures. That is, they contain sub-subatomic particles. These particles will not be covered in this textbook because it is not understood if or how they affect chemical behavior. As you will learn in coming chapters, behavior can be explained by considering only an atom's electrons, protons, and neutrons.

You should now have a solid understanding of the structure of a typical atom. But what makes an atom of one element different from an atom of another element? In the next section, you'll find out.

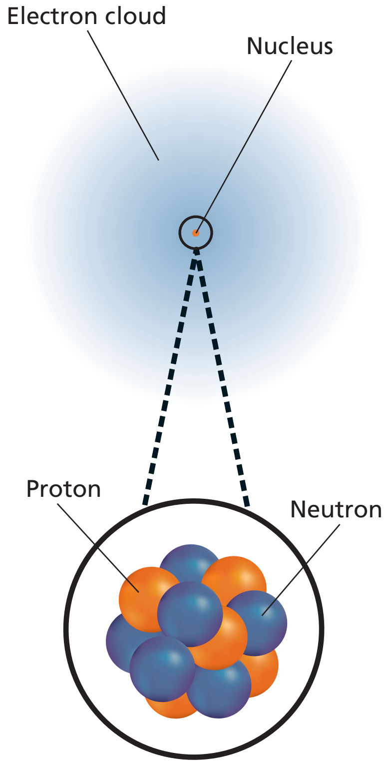


Figure 4-13

Atoms consist of a "cloud" of fast moving, negatively charged electrons surrounding a tiny, extremely dense nucleus containing positively charged protons and neutral neutrons. The nucleus contains virtually all of the atom's mass, but occupies only about one ten-thousandth the volume of the atom.

Section 4.2 Assessment

6. Briefly evaluate the experiments that led to the conclusion that electrons were negatively charged particles found in all matter.
7. Describe the structure of a typical atom. Be sure to identify where each subatomic particle is located.
8. Make a table comparing the relative charge and mass of each of the subatomic particles.
9. **Thinking Critically** Compare and contrast Thomson's plum pudding atomic model with Rutherford's nuclear atomic model.
10. **Graphing** Make a timeline graph of the development of modern atomic theory. Be sure to include the discovery of each subatomic particle.

Section 4.3 How Atoms Differ

Objectives

- **Explain** the role of atomic number in determining the identity of an atom.
- **Define** an isotope and **explain** why atomic masses are not whole numbers.
- **Calculate** the number of electrons, protons, and neutrons in an atom given its mass number and atomic number.

Vocabulary

atomic number
isotope
mass number
atomic mass unit (amu)
atomic mass

Look at the periodic table on the inside back cover of this textbook. As you can see, there are more than 110 different elements. This means that there are more than 110 different kinds of atoms. What makes an atom of one element different from an atom of another element? You know that all atoms are made up of electrons, protons, and neutrons. Thus, you might suspect that atoms somehow differ in the number of these particles. If so, you are correct.

Atomic Number

Not long after Rutherford's gold foil experiment, the English scientist Henry Moseley (1887–1915) discovered that atoms of each element contain a unique positive charge in their nuclei. Thus, the number of protons in an atom identifies it as an atom of a particular element. The number of protons in an atom is referred to as the element's **atomic number**. Look again at the periodic table and you will see that the atomic number determines the element's position in the table. Consider hydrogen, located at the top left of the table. The information provided by the periodic table for hydrogen is shown in **Figure 4-14**. Note that above the symbol for hydrogen (H), you see the number 1. This number, which corresponds to the number of protons in a hydrogen atom, is the atomic number of hydrogen. Hydrogen atoms always contain a single proton. Moving across the periodic table to the right, you'll next come to helium (He). Helium has an atomic number of 2, and thus has two protons in its nucleus. The next row begins with lithium (Li), atomic number 3, followed by beryllium (Be), atomic number 4, and so on. As you can see, the periodic table is organized left-to-right and top-to-bottom by increasing atomic number. How many protons does a gold atom contain? A silver atom?

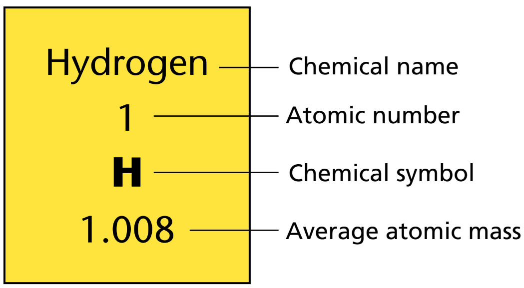
Remember that because all atoms are neutral, the number of protons and electrons in an atom must be equal. Thus, once you know the atomic number of an element, you know both the number of protons and the number of electrons an atom of that element contains.

$$\text{Atomic number} = \text{number of protons} = \text{number of electrons}$$

For instance, an atom of lithium, atomic number of 3, contains three protons and three electrons. How many electrons does an atom of element 97 contain?

Figure 4-14

The atomic number of an element equals the positive charge contained in its nucleus.



Hydrogen, with an atomic number of 1, is the first element in the periodic table. A hydrogen atom has one proton and a charge of 1+ in its nucleus.

EXAMPLE PROBLEM 4–1

Using Atomic Number

Complete the following table.

Composition of Several Elements

	Element	Atomic number	Protons	Electrons
a.	Pb	82	___	___
b.	___	___	8	___
c.	___	___	___	30

1. Analyze the Problem

You are given the information in the table. Apply the relationship among atomic number, number of protons, and number of electrons to complete most of the table. Once the atomic number is known, use the periodic table to identify the element.

2. Solve for the Unknown

Apply the atomic number relationship and then consult the periodic table to identify the element.

a. Atomic number = number of protons = number of electrons
82 = number of protons = number of electrons
Element 82 is lead (Pb).

b. Atomic number = number of protons = number of electrons
Atomic number = 8 = number of electrons
Element 8 is oxygen (O).

c. Atomic number = number of protons = number of electrons
Atomic number = number of protons = 30
Element 30 is zinc (Zn).

The completed table is shown below.

Composition of Several Elements

	Element	Atomic number	Protons	Electrons
a.	Pb	82	82	82
b.	O	8	8	8
c.	Zn	30	30	30

3. Evaluate the Answer

The answers agree with atomic numbers and element symbols given in the periodic table.

PRACTICE PROBLEMS

11. How many protons and electrons are in each of the following atoms?

a. boron
b. radon
c. platinum
d. magnesium

12. An atom of an element contains 66 electrons. What element is it?

13. An atom of an element contains 14 protons. What element is it?

For more practice with problems using atomic numbers, go to **Supplemental Practice Problems** in Appendix A.

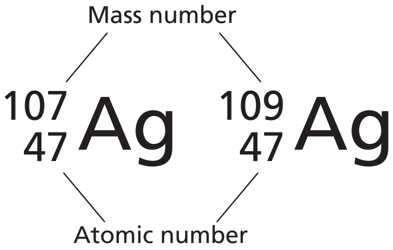


Figure 4-15

Ag is the chemical symbol for the silver used in these coins. The silver in each coin is comprised of 51.84% silver-107 ($^{107}_{47}Ag$) isotopes and 48.16% silver-109 ($^{109}_{47}Ag$) isotopes.

Isotopes and Mass Number

Earlier you learned that Dalton's atomic theory was wrong about atoms being indivisible. It was also incorrect in stating that all atoms of a particular element are identical. While it is true that all atoms of a particular element have the same number of protons and electrons, the number of neutrons on their nuclei may differ. For example, there are three different types of potassium atoms. All three types contain 19 protons (and thus 19 electrons). However, one type of potassium atom contains 20 neutrons, another contains 21 neutrons, and still another 22 neutrons. Atoms such as these, with the same number of protons but different numbers of neutrons, are called **isotopes**.

In nature most elements are found as a mixture of isotopes. Usually, no matter where a sample of an element is obtained, the relative abundance of each isotope is constant. For example, in a banana, which is a rich source of potassium, 93.25% of the potassium atoms have 20 neutrons, 6.7302% will have 22 neutrons, and a scant 0.0117% will have 21 neutrons. In another banana, or in a totally different source of potassium, the percentage composition of the potassium isotopes will still be the same.

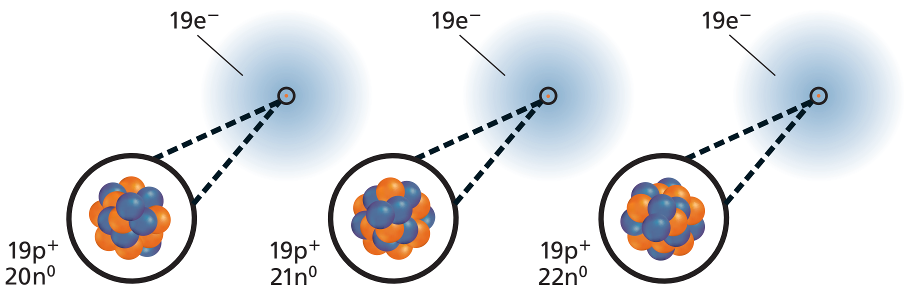
As you might expect, the isotopes do differ in mass. Isotopes containing more neutrons have a greater mass. In spite of differences in mass and the number of neutrons, isotopes of an atom have essentially the same chemical behavior. Why? Because, as you'll learn in greater detail later in this textbook, chemical behavior is determined by the number of electrons an atom has, not by its number of neutrons and protons. To make it easy to identify each of the various isotopes of an element, chemists add a number after the element's name. The number that is added is called the **mass number**, and it represents the sum of the number of protons and neutrons in the nucleus. For example, the potassium isotope with 19 protons and 20 neutrons has a mass number of 39 ($19 + 20 = 39$), and the isotope is called potassium-39. The potassium isotope with 19 protons and 21 neutrons has a mass number of 40 ($19 + 21 = 40$), and is called potassium-40. What is the mass number and name of the potassium isotope with 19 protons and 22 neutrons?

Chemists often write out isotopes using a shortened type of notation involving the chemical symbol, atomic number, and mass number, as shown in **Figure 4-15**. Note that the mass number is written as a superscript to the left of the chemical symbol, and the atomic number is written as a subscript to the left of the chemical symbol. The three potassium isotopes you have just learned

Figure 4-16

The three naturally occurring potassium isotopes are potassium-39, potassium-40, and potassium-41. How do their masses compare? Their chemical properties?

	Potassium-39	Potassium-40	Potassium-41
Protons	19	19	19
Neutrons	20	21	22
Electrons	19	19	19



CONTENTS

about are summarized in **Figure 4-16**. The number of neutrons in an isotope can be calculated from the atomic number and mass number.

Number of neutrons = mass number − atomic number

EXAMPLE PROBLEM 4–2

Using Atomic Number and Mass Number

A chemistry laboratory has analyzed the composition of isotopes of several elements. The composition data is given in the table at the right. Data for one of neon's three isotopes is given in the table. Determine the number of protons, electrons, and neutrons in the isotope of neon. Name the isotope and give its symbol.

Isotope Composition Data

Element	Atomic number	Mass number
a. Neon	10	22
b. Calcium	20	46
c. Oxygen	8	17
d. Iron	26	57
e. Zinc	30	64
f. Mercury	80	204

1. Analyze the Problem

You are given some data for neon in the table. The symbol for neon can be found from the periodic table. From the atomic number, the number of protons and electrons in the isotope are known. The number of neutrons in the isotope can be found by subtracting the atomic number from the mass number.

Known	Unknown
Element: neon	Number of protons, electrons, and neutrons = ?
Atomic number = 10	Name of isotope = ?
Mass number = 22	Symbol for isotope = ?

2. Solve for the Unknown

The number of protons equals the number of electrons which equals the atomic number.

Number of protons = number of electrons = atomic number = 10

Use the atomic number and the mass number to calculate the number of neutrons.

Number of neutrons = mass number − atomic number

Number of neutrons = 22 − 10 = 12

Use the element name and mass number to write the isotope's name.

neon-22

Use the chemical symbol, mass number, and atomic number to write out the isotope in symbolic notation form.

$^{22}_{10}\text{Ne}$

3. Evaluate the Answer

The relationships among number of electrons, protons, and neutrons have been applied correctly. The isotope's name and symbol are in the correct format.

PRACTICE PROBLEM

14. Determine the number of protons, electrons, and neutrons for isotopes b. through f. in the table above. Name each isotope, and write its symbol.

Practice!

For more practice with problems using atomic number and mass number, go to **Supplemental Practice Problems** in Appendix A.

Mass of Individual Atoms

Recall from **Table 4-1** that the masses of both protons and neutrons are approximately 1.67×10^{-24} g. While this is a very small mass, the mass of an electron is even smaller—only about $\frac{1}{1840}$ that of a proton or neutron. Because these extremely small masses expressed in scientific notation are difficult to work with, chemists have developed a method of measuring the mass of an atom relative to the mass of a specifically chosen atomic standard. That standard is the carbon-12 atom. Scientists assigned the carbon-12 atom a mass of exactly 12 atomic mass units. Thus, one **atomic mass unit (amu)** is defined as $\frac{1}{12}$ the mass of a carbon-12 atom. Although a mass of 1 amu is very nearly equal to the mass of a single proton or a single neutron, it is important to realize that the values are slightly different. As a result, the mass of silicon-30, for example, is 29.974 amu, and not 30 amu. **Table 4-2** gives the masses of the subatomic particles in terms of amu.

Table 4-2

Masses of Subatomic Particles

Particle	Mass (amu)
Electron	0.000 549
Proton	1.007 276
Neutron	1.008 665

Because an atom's mass depends mainly on the number of protons and neutrons it contains, and because protons and neutrons have masses close to 1 amu, you might expect the atomic mass of an element to always be very near a whole number. This, however, is often not the case. The explanation involves how atomic mass is defined. The **atomic mass** of an element is the weighted

miniLAB

Modeling Isotopes

Formulating Models Because they have different compositions, pre- and post-1982 pennies can be used to model an element with two naturally occurring isotopes. From the penny "isotope" data, the mass of each penny isotope and the average mass of a penny can be determined.

Materials bag of pre- and post-1982 pennies, balance

Procedure

1. Get a bag of pennies from your teacher, and sort the pennies by date into two groups: pre-1982 pennies and post-1982 pennies. Count and record the total number of pennies and the number of pennies in each group.
2. Use the balance to determine the mass of ten pennies from each group. Record each mass to the nearest 0.01 g. Divide the total mass of each group by ten to get the average mass of a pre- and post-1982 penny "isotope."

Analysis

1. Using data from step 1, calculate the percentage abundance of each group. To do this, divide the number of pennies in each group by the total number of pennies.
2. Using the percentage abundance of each "isotope" and data from step 2, calculate the atomic mass of a penny. To do this, use the following equation for each "isotope."

   mass contribution = (% abundance)(mass)

   Sum the mass contributions to determine the atomic mass.
3. Would the atomic mass be different if you received another bag of pennies containing a different mixture of pre- and post-1982 pennies? Explain.
4. In step 2, instead of measuring and using the mass of a single penny of each group, the average mass of each type of penny was determined. Explain why.

Calculating the Weighted Average Atomic Mass of Chlorine

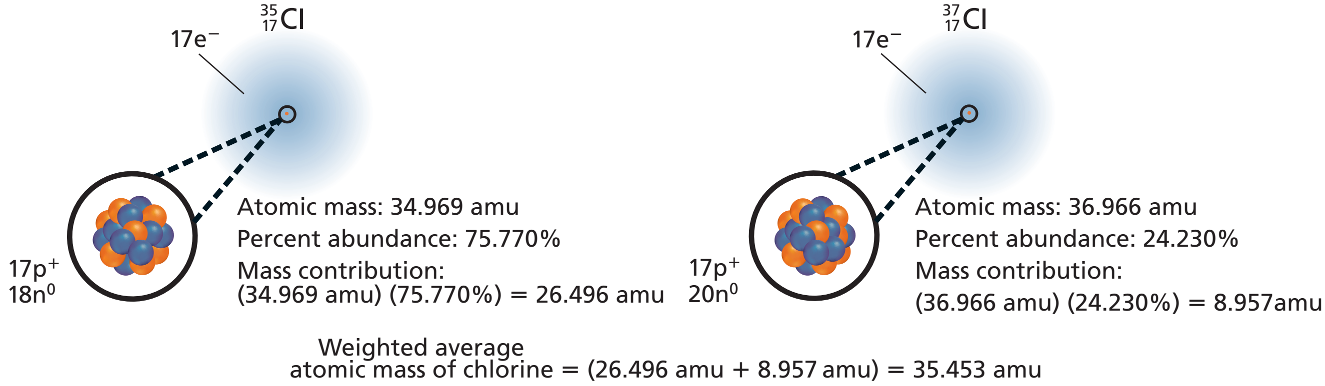


Figure 4-17

To determine the weighted average atomic mass of chlorine, the mass contribution of each of the two isotopes is calculated, and then those two values are added together.

average mass of the isotopes of that element. For example, the atomic mass of chlorine is 35.453 amu. Chlorine exists naturally as a mixture of about 75% chlorine-35 and 25% chlorine-37. Because atomic mass is a weighted average, the chlorine-35 atoms, which exist in greater abundance than the chlorine-37 atoms, have a greater effect in determining the atomic mass. The atomic mass of chlorine is calculated by summing the products of each isotope's percent abundance times its atomic mass. See **Figure 4-17**. For hands-on practice in calculating atomic mass, do the **miniLAB** on the previous page.

You can calculate the atomic mass of any element if you know its number of naturally occurring isotopes, their masses, and their percent abundances. The following Example Problem and Practice Problems will provide practice in calculating atomic mass.

EXAMPLE PROBLEM 4–3

Calculating Atomic Mass

Given the data in the table at the right, calculate the atomic mass of unknown element *X*. Then, identify the unknown element, which is used medically to treat some mental disorders.

Isotope Abundance for Element *X*

Isotope	Mass (amu)	Percent abundance
^6X	6.015	7.5%
^7X	7.016	92.5%

1. Analyze the Problem

You are given the data in the table. Calculate the atomic mass by multiplying the mass of each isotope by its percent abundance and summing the results. Use the periodic table to confirm the calculation and identify the element.

Known

For isotope ^6X:
- mass = 6.015 amu
- abundance = 7.50% = 0.0750

For isotope ^7X:
- mass = 7.016 amu
- abundance = 92.5% = 0.925

Unknown

atomic mass of X = ? amu
name of element X = ?

2. Solve for the Unknown

Calculate each isotope's contribution to the atomic mass.

For ^6X: Mass contribution = (mass)(percent abundance)

mass contribution = (6.015 amu)(0.0750) = 0.451 amu

Continued on next page

For ^7X: Mass contribution = (mass)(percent abundance)

mass contribution = (7.016 amu)(0.925) = 6.490 amu

Sum the mass contributions to find the atomic mass.

Atomic mass of X = (0.451 amu + 6.490 amu) = 6.941 amu

Use the periodic table to identify the element.

The element with a mass of 6.941 amu is lithium (Li).

3. Evaluate the Answer

The result of the calculation agrees with the atomic mass given in the periodic table. The masses of the isotopes have four significant figures, so the atomic mass is also expressed with four significant figures.

PRACTICE PROBLEMS

Practice!

For more practice with atomic mass problems, go to **Supplemental Practice Problems** in Appendix A.

15. Boron has two naturally occurring isotopes: boron-10 (abundance = 19.8%, mass = 10.013 amu), boron-11 (abundance = 80.2%, mass = 11.009 amu). Calculate the atomic mass of boron.

16. Helium has two naturally occurring isotopes, helium-3 and helium-4. The atomic mass of helium is 4.003 amu. Which isotope is more abundant in nature? Explain.

17. Calculate the atomic mass of magnesium. The three magnesium isotopes have atomic masses and relative abundances of 23.985 amu (78.99%), 24.986 amu (10.00%), and 25.982 amu (11.01%).

Analyzing an element's mass can give you insight into what the most abundant isotope for the element may be. For example, note that fluorine (F) has an atomic mass that is extremely close to a value of 19 amu. If fluorine had several fairly abundant isotopes, it would be unlikely that its atomic mass would be so close to a whole number. Thus, you might conclude that virtually all naturally occurring fluorine is probably in the form of fluorine-19 ($^{19}_{9}F$). You would be correct, as 100% of naturally occurring fluorine is in the form of fluorine-19. While this type of reasoning generally works well, it is not foolproof. Consider bromine (Br), with an atomic mass of 79.904 amu. With a mass so close to 80 amu, it seems likely that the most common bromine isotope would be bromine-80 ($^{80}_{35}Br$). This is not the case, however. Bromine's two isotopes, bromine-79 (78.918 amu, 50.69%) and bromine-81 (80.917 amu, 49.31%), have a weighted average atomic mass of approximately 80 amu, but there is no bromine-80 isotope.

Section 4.3 Assessment

18. Which subatomic particle identifies an atom as that of a particular element? How is this particle related to the atom's atomic number?

19. What is an isotope? Give an example of an element with isotopes.

20. Explain how the existence of isotopes is related to atomic masses not being whole numbers.

21. **Thinking Critically** Nitrogen has two naturally occurring isotopes, N-14 and N-15. The atomic mass of nitrogen is 14.007 amu. Which isotope is more abundant in nature? Explain.

22. **Communicating** List the steps in the process of calculating average atomic mass given data about the isotopes of an element.

Section 4.4

Unstable Nuclei and Radioactive Decay

You now have a good understanding of the basic structure of matter and how matter interacts and changes through processes called chemical reactions. With the information you have just learned about the atom's nuclear nature, you are ready to learn about a very different type of reaction—the nuclear reaction. This section introduces you to some of the changes that can take place in a nucleus; you will revisit and further explore this topic in Chapter 25 when you study nuclear chemistry.

Objectives

- **Explain** the relationship between unstable nuclei and radioactive decay.
- **Characterize** alpha, beta, and gamma radiation in terms of mass and charge.

Vocabulary

nuclear reaction
radioactivity
radiation
radioactive decay
alpha radiation
alpha particle
nuclear equation
beta radiation
beta particle
gamma ray

Radioactivity

Recall from Chapter 3 that a chemical reaction involves the change of one or more substances into new substances. Although atoms may be rearranged, their identities do not change during the reaction. You may be wondering why atoms of one element do not change into atoms of another element during a chemical reaction. The reason has to do with the fact that chemical reactions involve only an atom's electrons—the nucleus remains unchanged.

As you learned in the previous section, the number of protons in the nucleus determines the identity of an atom. Thus, because there are no changes in the nuclei during a chemical reaction, the identities of the atoms do not change. There are, however, reactions that do involve an atom of one element changing into an atom of another element. These reactions, which involve a change in an atom's nucleus, are called **nuclear reactions**.

In the late 1890s, scientists noticed that some substances spontaneously emitted radiation in a process they called **radioactivity**. The rays and particles emitted by the radioactive material were called **radiation**. Scientists studying radioactivity soon made an important discovery—radioactive atoms undergo significant changes that can alter their identities. In other words, by emitting radiation, atoms of one element can change into atoms of another element. This discovery was a major breakthrough, as no chemical reaction had ever resulted in the formation of new kinds of atoms.

Radioactive atoms emit radiation because their nuclei are unstable. Unstable systems, whether they're atoms or the pencil standing on its sharpened tip shown in **Figure 4-18a**, gain stability by losing energy. As you can see in **Figure 4-18b** and **Figure 4-18c**, the pencil gains stability (and loses energy) by toppling over. When resting flat on the table top, the pencil has

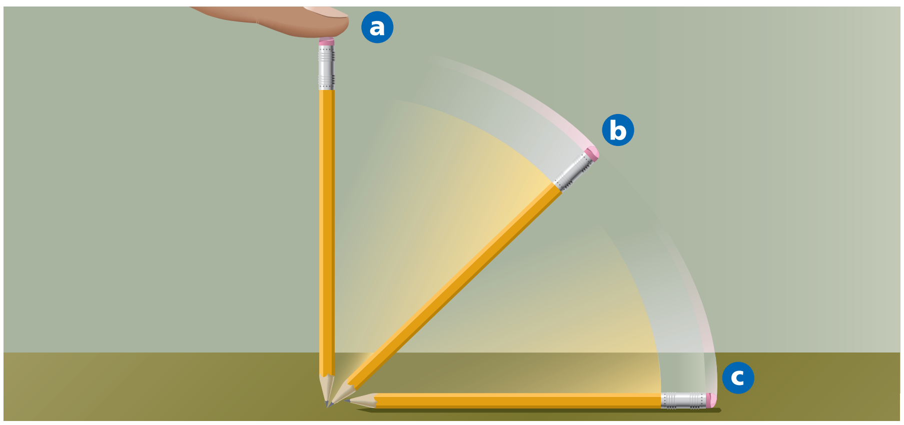


Figure 4-18

Unstable systems, such as this pencil momentarily standing on its tip, gain stability by losing energy. In this case, the pencil loses gravitational potential energy as it topples over. Unstable atoms also gain stability by losing energy—they lose energy by emitting radiation.

less gravitational potential energy than it did in its upright position, and thus, it is more stable. Unstable nuclei lose energy by emitting radiation in a spontaneous process (a process that does not require energy) called **radioactive decay**. Unstable radioactive atoms undergo radioactive decay until they form stable nonradioactive atoms, often of a different element. Several types of radiation are commonly emitted during radioactive decay.

Types of Radiation

Scientists began researching radioactivity in the late 1800s. By directing radiation from a radioactive source between two electrically charged plates, scientists were able to identify three different types of radiation. As you can see in **Figure 4-19**, some of the radiation was deflected toward the negatively charged plate, some was deflected toward the positively charged plate, and some was not deflected at all.

Alpha radiation Scientists named the radiation that was deflected toward the negatively charged plate **alpha radiation**. This radiation is made up of **alpha particles**. Each alpha particle contains two protons and two neutrons, and thus has a 2+ charge. As you know, opposite electrical charges attract. So the 2+ charge explains why alpha particles are attracted to the negatively charged plate shown in **Figure 4-19**. An alpha particle is equivalent to a helium-4 nucleus and is represented by $^{4}_{2}He$ or α. The alpha decay of radioactive radium-226 into radon-222 is shown below.

$$\underset{\text{radium-226}}{^{226}_{88}Ra} \longrightarrow \underset{\text{radon-222}}{^{222}_{86}Rn} + \underset{\text{alpha particle}}{^{4}_{2}He}$$

Note also that a new element, radon (Rn), is created as a result of the alpha decay of the unstable radium-226 nucleus. The type of equation shown above is known as a **nuclear equation** because it shows the atomic number and mass number of the particles involved. It is important to note that both mass number and atomic number are conserved in nuclear equations. The accounting of atomic numbers and mass numbers below shows that they are conserved.

$$^{226}_{88}Ra \longrightarrow {^{222}_{86}Rn} + {^{4}_{2}He}$$

$$\text{Atomic number: } 88 \longrightarrow 86 + 2$$

$$\text{Mass number: } 226 \longrightarrow 222 + 4$$

Careers Using Chemistry

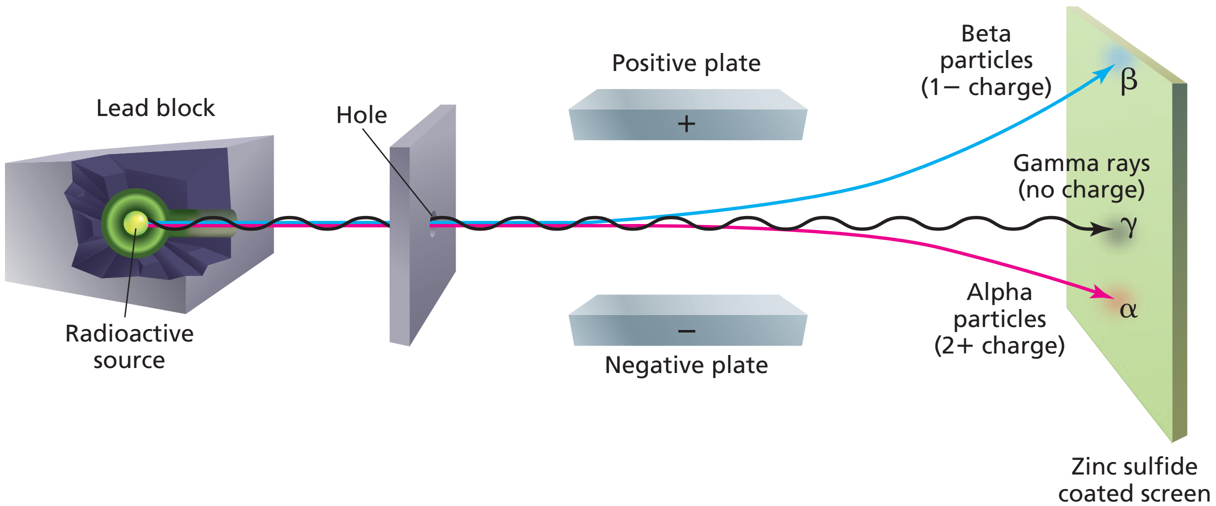
Radiation Protection Technician

Do you like the idea of protecting the health of others? Would you enjoy the challenge of removing contaminated materials? If so, consider a career as a radiation protection technician.

Radiation protection technicians use radiation measuring instruments to locate and assess the risk posed by contaminated materials. Then they decontaminate the area with high-pressure cleaning equipment and remove the radioactive materials. The work often requires the technicians to wear protective suits. The physically demanding work is carefully planned and carried out, with an emphasis on safety.

Figure 4-19

Because alpha, beta, and gamma radiation possess different amounts of electrical charge, they are affected differently by an electric field. Gamma rays, which carry no charge, are not deflected by the electric field.



Beta radiation Scientists named the radiation that was deflected toward the positively charged plate **beta radiation**. This radiation consists of fast moving electrons called **beta particles**. Each beta particle is an electron with a 1– charge. The negative charge of the beta particle explains why it is attracted to the positively charged plate shown in **Figure 4-19**. Beta particles are represented by the symbol ${}^{0}_{-1}\beta$. The beta decay of radioactive carbon-14 into nitrogen-14 is shown below.

$$\underset{\text{carbon-14}}{{}^{14}_{6}\text{C}} \rightarrow \underset{\text{nitrogen-14}}{{}^{14}_{7}\text{N}} + \underset{\text{beta particle}}{{}^{0}_{-1}\beta}$$

The beta decay of unstable carbon-14 results in the creation of the new atom nitrogen (N).

Gamma radiation The third common type of radiation is called gamma radiation, or gamma rays. **Gamma rays** are high-energy radiation that possess no mass and are denoted by the symbol ${}^{0}_{0}\gamma$. Because they possess no electrical charge, gamma rays are not deflected by electric or magnetic fields. Gamma rays usually accompany alpha and beta radiation, and they account for most of the energy lost during the radioactive decay process. For example, gamma rays accompany the alpha decay of uranium-238.

$$\underset{\text{uranium-238}}{{}^{238}_{92}\text{U}} \rightarrow \underset{\text{thorium-234}}{{}^{234}_{90}\text{Th}} + \underset{\text{alpha particle}}{{}^{4}_{2}\text{He}} + \underset{\text{gamma rays}}{2\,{}^{0}_{0}\gamma}$$

Because gamma rays are massless, the emission of gamma rays by themselves cannot result in the formation of a new atom. **Table 4-3** summarizes the basic characteristics of alpha, beta, and gamma radiation.

Table 4-3

Characteristics of Alpha, Beta, and Gamma Radiation

Radiation type	Symbol	Mass (amu)	Charge
Alpha	${}^{4}_{2}\text{He}$	4	2+
Beta	${}^{0}_{-1}\beta$	$\frac{1}{1840}$	1–
Gamma	${}^{0}_{0}\gamma$	0	0

Nuclear stability You may be wondering why some atoms are stable while others are not. The primary factor in determining an atom's stability is its ratio of neutrons to protons. The details of how the neutron-to-proton ratio determines stability will be covered in Chapter 25. For now, it is enough that you know that atoms containing either too many or too few neutrons are unstable. Unstable nuclei lose energy through radioactive decay in order to form a nucleus with a stable composition of neutrons and protons. Alpha and beta particles are emitted during radioactive decay, and these emissions affect the neutron-to-proton ratio of the newly created nucleus. Eventually, radioactive atoms undergo enough radioactive decay to form stable, nonradioactive atoms. This explains why there are so few radioactive atoms found in nature—most of them have already decayed into stable atoms.

Section 4.4 Assessment

23. Explain how unstable atoms gain stability. What determines whether or not an atom is stable?

24. Create a table comparing the mass and charge of alpha, beta, and gamma radiation.

25. In writing a balanced nuclear equation, what must be conserved?

26. **Thinking Critically** Explain how a nuclear reaction differs from a chemical reaction.

27. **Classifying** Classify each of the following as a chemical reaction, a nuclear reaction, or neither.

a. Thorium emits a beta particle.

b. Two atoms share electrons to form a bond.

c. A sample of pure sulfur emits heat energy as it slowly cools.

d. A piece of iron rusts.

Very Small Particles

This laboratory investigation will help you conceptualize the size of an atom. You will experiment with a latex balloon containing a vanilla bean extract. Latex is a polymer, meaning that it is a large molecule (a group of atoms that act as a unit) that is made up of a repeating pattern of smaller molecules. The scent of the vanilla extract will allow you to trace the movement of its molecules through the walls of the solid latex balloon.

Problem

How small are the atoms that make up the molecules of the balloon and the vanilla extract? How can you conclude the vanilla molecules are in motion?

Objectives

- **Observe** the movement of vanilla molecules based on detecting their scent.
- **Infer** what the presence of the vanilla scent means in terms of the size and movement of its molecules.
- **Formulate models** that explain how small molecules in motion can pass through an apparent solid.
- **Hypothesize** about the size of atoms that make up matter.

Materials

vanilla extract or flavoring
9-inch latex balloon (2)
dropper

Safety Precautions

- **Always wear safety goggles and a lab apron.**
- **Be careful not to cut yourself when using a sharp object to deflate the balloon.**

Pre-Lab

1. Read the entire CHEMLAB.
2. Describe a polymer and give an example.
3. Identify constants in the experiment.
4. What is the purpose of the vanilla extract?
5. As a liquid evaporates, predict what you think will happen to the temperature of the remaining liquid.
6. When you smell an aroma, is your nose detecting a particle in the solid, liquid, or gas phase?
7. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table in which to record your data and observations.

Data Table

Observations		Initial	Final
Balloon 1 with vanilla	Relative size		
	Relative temperature		
Balloon 2 without vanilla	Relative size		
	Relative temperature		

Procedure

1. Using the medicine dropper, add 25 to 30 drops of vanilla extract to the first balloon.

2. Inflate the balloon so its walls are tightly stretched, but not stretched so tightly that the balloon is in danger of bursting. Try to keep the vanilla in one location as the balloon is inflated. Tie the balloon closed.
3. Feel the outside of the balloon where the vanilla is located and note the temperature of this area relative to the rest of the balloon. Record your observations in the data table.
4. Use only air to inflate a second balloon to approximately the same size as that of the first, and tie it closed. Feel the outside of the second balloon. Make a relative temperature comparison to that of the first balloon. Record your initial observations.
5. Place the inflated balloons in a small, enclosed area such as a closet or student locker.
6. The next day, repeat the observations in steps 3 and 4 after the vanilla has dried inside the balloon. Record these final observations.
7. To avoid splattering your clothes with dark brown vanilla, do not deflate the balloon until the vanilla has dried inside.

Cleanup and Disposal

1. After the vanilla has dried, deflate the balloon by puncturing it with a sharp object.
2. Dispose of the pieces of the balloon as directed by your teacher.

Analyze and Conclude

1. **Observing and Inferring** How did the relative volumes of balloons 1 and 2 change after 24 hours? Explain.
2. **Observing and Inferring** By comparing the relative temperatures of balloons 1 and 2, what can you conclude about the temperature change as the vanilla evaporated? Explain.
3. **Observing and Inferring** Did the vanilla's odor get outside the balloon and fill the enclosed space? Explain.
4. **Predicting** Do you think vanilla will leak more rapidly from a fully inflated balloon or from a half-inflated balloon? Explain.
5. **Hypothesizing** Write a hypothesis that explains your observations.
6. **Comparing and Contrasting** Compare your hypothesis to Dalton's atomic theory. In what ways is it similar? How is it different?
7. **Error Analysis** What factors might affect the results of different groups that performed the experiment? What types of errors might have occurred during the procedure?

Real-World Chemistry

1. Explain why helium-filled, Mylar-foil balloons can float freely for several weeks, but latex balloons for less than 24 hours.
2. How are high-pressure gases stored for laboratory and industrial use to prevent loss?

CHEMISTRY and Society

Nanotechnology

Imagine a technology able to produce roads that repave themselves, greenhouses productive enough to end starvation, and computers the size of cells. Sound like science fiction? It may not be.

Starting at the bottom

This yet-to-be realized technology is generally known as nanotechnology. The prefix *nano-* means one billionth. A nanometer is roughly the size of several atoms put together. The goal of nanotechnology is to manipulate individual atoms in order to create a wide variety of products.

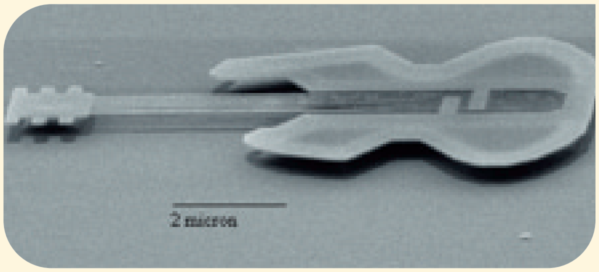
In order to manipulate the immense number of atoms required to make a product, scientists plan on constructing tiny robots called nanorobots. Nanorobots would have two objectives—to manipulate atoms and to copy themselves (self-replication). Through self-replication, countless nanorobots could be created. This work force of nanorobots would then work together to quickly and efficiently assemble new products.

Possibilities and effects

The benefits of nanotechnology could potentially go beyond those of all other existing technologies. The quality and reliability of manufactured products could improve dramatically. For example, a brick could repair itself after cracks form, and a damaged road could repave itself. Furthermore, even as the quality and capability of products increase, their prices would decrease. With the use of nanorobot workforces and a readily available supply of atoms, the cost of atom-assembled products would be low.

Unlike many technological advances of the past, nanotechnology could also benefit the environment. The need for traditional raw materials, such as trees and coal, would be greatly reduced. Also, because the atoms are arranged individually, the amount of waste could be carefully controlled and limited. Thin materials called nanomembranes may even be able to filter existing pollutants out of air and water.

Advances in medicine could also be amazing. Nanodevices smaller than human cells could be used to detect health problems, repair cells, and carry medicines to specific sites in the body.



When will it happen?

The promise of nanotechnology is still years away. Some think nanodevices will be available in the next decade, whereas others expect the technology to take much longer to develop. So far, researchers have accomplished rearranging atoms into specific shapes such as letters and symbols, and have also succeeded in developing several very simple nanodevices. One such device, a nanoguitar, made of crystalline silicon, is about 1/20th the width of a single human hair.

Investigating the Issue

1. **Communicating Ideas** Read about the Industrial Revolution and write a brief essay describing how technological advances affected society.

2. **Debating the Issue** Nanotechnology supporters argue that dangers posed by the technology will be addressed as nanotechnology is developed. Should researchers be prevented from developing nanotechnology?

CHEMISTRY Online CLICK HERE

Visit the Chemistry Web site at **science.glencoe.com** to find links to more information about nanotechnology and the issues that surround it.

Summary

4.1 Early Theories of Matter

- The Greek philosopher Democritus was the first person to propose the existence of atoms.
- In 1808, Dalton proposed his atomic theory, which was based on numerous scientific experiments.
- All matter is composed of atoms. An atom is the smallest particle of an element that maintains the properties of that element. Atoms of one element are different from atoms of other elements.

4.2 Subatomic Particles and the Nuclear Atom

- Atoms are composed of negatively charged electrons, neutral neutrons, and positively charged protons. Electrons have a 1− charge, protons have a 1+ charge, and neutrons have no charge. Both protons and electrons have masses approximately 1840 times that of an electron.
- The nucleus of an atom contains all of its positive charge and nearly all of its mass.
- The nucleus occupies an extremely small volume of space at the center of an atom. Most of an atom consists of empty space surrounding the nucleus through which the electrons move.

4.3 How Atoms Differ

- The number of protons in an atom uniquely identifies an atom. This number of protons is the atomic number of the atom.
- Atoms have equal numbers of protons and electrons, and thus, no overall electrical charge.
- An atom's mass number is equal to its total number of protons and neutrons.
- Atoms of the same element with different numbers of neutrons and different masses are called isotopes.
- The atomic mass of an element is a weighted average of the masses of all the naturally occurring isotopes of that element.

4.4 Unstable Nuclei and Radioactive Decay

- Chemical reactions involve changes in the electrons surrounding an atom. Nuclear reactions involve changes in the nucleus of an atom.
- The neutron-to-proton ratio of an atom's nucleus determines its stability. Unstable nuclei undergo radioactive decay, emitting radiation in the process.
- Alpha particles are equivalent to the nuclei of helium atoms, and are represented by ${}^{4}_{2}He$ or α. Alpha particles have a charge of 2+.
- Beta particles are high-speed electrons and are represented by ${}^{0}_{-1}\beta$. Beta particles have a 1− charge.
- Gamma rays are high-energy radiation and are represented by the symbol ${}^{0}_{0}\gamma$. Gamma rays have no electrical charge and no mass.

Key Equations and Relationships

- Determining the number of protons and electrons

$$\text{Atomic number} = \text{number of protons} = \text{number of electrons} \quad \text{(p. 98)}$$

- Determining the number of neutrons

$$\text{Number of neutrons} = \text{mass number} - \text{atomic number} \quad \text{(p. 101)}$$

Vocabulary

- alpha particle (p. 106)
- alpha radiation (p. 106)
- atom (p. 90)
- atomic mass (p. 102)
- atomic mass unit (amu) (p. 102)
- atomic number (p. 98)
- beta particle (p. 107)
- beta radiation (p. 107)
- cathode ray (p. 92)
- Dalton's atomic theory (p. 89)
- electron (p. 93)
- gamma ray (p. 107)
- isotope (p. 100)
- mass number (p. 100)
- neutron (p. 96)
- nuclear equation (p. 106)
- nuclear reaction (p. 105)
- nucleus (p. 95)
- proton (p. 96)
- radiation (p. 105)
- radioactive decay (p. 106)
- radioactivity (p. 105)

CHAPTER 4 ASSESSMENT

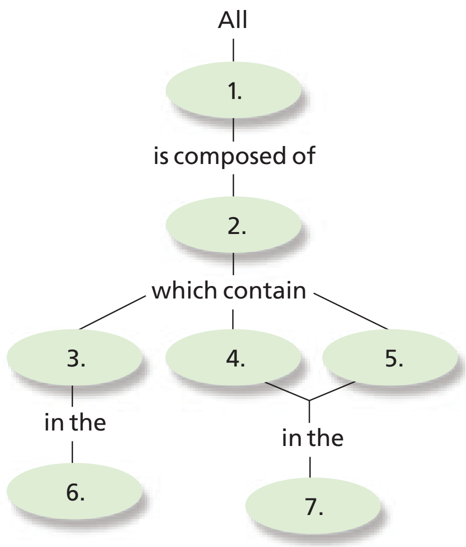
CHEMISTRY Online

CLICK HERE

Go to the Chemistry Web site at **science.glencoe.com** *or use the* **Chemistry CD-ROM** *for additional Chapter 4 Assessment.*

Concept Mapping

28. Complete the concept map using the following terms: electrons, matter, neutrons, nucleus, empty space around nucleus, protons, and atoms.



Mastering Concepts

29. Who originally proposed the concept that matter was composed of tiny indivisible particles? (4.1)

30. Whose work is credited with being the beginning of modern atomic theory? (4.1)

31. Explain why Democritus was unable to experimentally verify his ideas. (4.1)

32. State the main points of Dalton's atomic theory using your own words. Which parts of Dalton's theory were later found to be in error? Explain why. (4.1)

33. Explain how Dalton's atomic theory offered a convincing explanation of the observation that mass is conserved in chemical reactions. (4.1)

34. Which subatomic particle was discovered by researchers working with cathode ray tubes? (4.2)

35. What experimental results led to the conclusion that electrons were part of all forms of matter? (4.2)

36. What is the charge and mass of a single electron? (4.2)

37. List the strengths and weaknesses of Rutherford's nuclear model of the atom. (4.2)

38. What particles are found in the nucleus of an atom? What is the net charge of the nucleus? (4.2)

39. Explain what keeps the electrons confined in the space surrounding the nucleus. (4.2)

40. Describe the flow of a cathode ray inside a cathode ray tube. (4.2)

41. Which outdated atomic model could be likened to chocolate chip cookie dough? (4.2)

42. What caused the deflection of the alpha particles in Rutherford's gold foil experiment? (4.2)

43. Which subatomic particles account for most all of an atom's mass? (4.2)

44. How is an atom's atomic number related to its number of protons? To its number of electrons? (4.2)

45. What is the charge of the nucleus of element 89? (4.2)

46. Explain why atoms are electrically neutral. (4.2)

47. Does the existence of isotopes contradict part of Dalton's original atomic theory? Explain. (4.3)

48. How do isotopes of a given element differ? How are they similar? (4.3)

49. How is the mass number related to the number of protons and neutrons an atom has? (4.3)

50. What do the superscript and subscript in the notation $^{40}_{19}K$ represent? (4.3)

51. Explain how to determine the number of neutrons an atom contains if you know its mass number and its atomic number. (4.3)

52. Define the atomic mass unit. What were the benefits of developing the atomic mass unit as a standard unit of mass? (4.3)

53. What type of reaction involves changes in the nucleus of an atom? (4.4)

54. Explain how energy loss and nuclear stability are related to radioactive decay. (4.4)

55. Explain what must occur before a radioactive atom ceases to undergo further radioactive decay. (4.4)

56. Write the symbols used to denote alpha, beta, and gamma radiation and give their mass and charge. (4.4)

57. What change in mass number occurs when a radioactive atom emits an alpha particle? A beta particle? A gamma particle? (4.4)

58. What is the primary factor determining whether or not an atom is stable or unstable? (4.4)

Mastering Problems

Atomic Number and Mass Number (4.3)

59. How many protons and electrons are contained in an atom of element 44?

60. For each of the following chemical symbols, determine the element name and the number of protons and electrons an atom contains.

a. V
b. Mn
c. Ir
d. S

61. A carbon atom has a mass number of 12 and an atomic number of 6. How many neutrons does it have?

62. An isotope of mercury has 80 protons and 120 neutrons. What is the mass number of this isotope?

63. An isotope of xenon has an atomic number of 54 and contains 77 neutrons. What is the xenon isotope's mass number?

64. How many electrons, protons, and neutrons are contained in each of the following atoms?

a. $^{132}_{55}Cs$
b. $^{59}_{27}Co$
c. $^{163}_{69}Tm$
d. $^{70}_{30}Zn$

65. How many electrons, protons, and neutrons are contained in each of the following atoms?

a. gallium-64
b. fluorine-23
c. titanium-48
d. helium-8

Atomic Mass (4.3)

66. Chlorine, which has an atomic mass of 35.453 amu, has two naturally occurring isotopes, Cl-35 and Cl-37. Which isotope occurs in greater abundance? Explain.

67. Silver has two isotopes, $^{107}_{47}Ag$ has a mass of 106.905 amu (52.00%), and $^{109}_{47}Ag$ has a mass of 108.905 amu (48.00%). What is the atomic mass of silver?

68. Data for chromium's four naturally occuring isotopes is provided in the table below. Calculate chromium's atomic mass.

Chromium Isotope Data

Isotope	Percent abundance	Mass (amu)
Cr-50	4.35%	49.946
Cr-52	83.79%	51.941
Cr-53	9.50%	52.941
Cr-54	2.36%	53.939

Mixed Review

Sharpen your problem-solving skills by answering the following.

69. Describe a cathode ray tube and how it operates.

70. Explain how J. J. Thomson's determination of the charge-to-mass ratio of the electron led to the conclusion that atoms were composed of subatomic particles.

71. How did the actual results of Rutherford's gold foil experiment differ from the results he expected?

72. Complete the table below.

Composition of Various Isotopes

Isotope	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
		32	16		
				24	20
Zn-64					
	9			10	
	11	23			

73. Approximately how many times greater is the diameter of an atom than the diameter of its nucleus? Knowing that most of an atom's mass is contained in the nucleus, what can you conclude about the density of the nucleus?

74. Is the charge of a nucleus positive, negative, or zero? The charge of an atom?

75. Why are electrons in a cathode ray tube deflected by magnetic and electric fields?

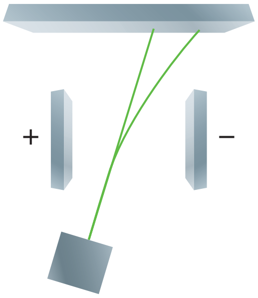
76. What was Henry Moseley's contribution to our understanding of the atom?

77. What is the mass number of potassium-39? What is the isotope's charge?

78. Boron-10 and boron-11 are the naturally occurring isotopes of elemental boron. If boron has an atomic mass of 10.81 amu, which isotope occurs in greater abundance?

79. Calculate the atomic mass of titanium. The five titanium isotopes have atomic masses and relative abundances of 45.953 amu (8.00%), 46.952 amu (7.30%), 47.948 amu (73.80%), 48.948 amu (5.50%), and 49.945 amu (5.40%).

80. Identify the two types of radiation shown in the figure below. Explain your reasoning.



81. Describe how each type of radiation affects an atom's atomic number and mass number.

82. Silicon is very important to the semiconductor manufacturing industry. The three naturally occurring isotopes of silicon are silicon-28, silicon-29, and silicon-30. Write the symbol for each.

Thinking Critically

83. Applying Concepts Which is greater, the number of compounds or the number of elements? The number of atoms or the number of isotopes? Explain.

84. Analyzing Information An element has three naturally occurring isotopes. What other information must you know in order to calculate the element's atomic mass?

85. Applying Concepts If atoms are primarily composed of empty space, why can't you pass your hand through a solid object?

86. Formulating Models Sketch a modern atomic model of a typical atom and identify where each type of subatomic particle would be located.

87. Applying Concepts Copper has two naturally occurring isotopes and an atomic mass of 63.546 amu. Cu-63 has a mass of 62.940 amu and an abundance of 69.17%. What is the identity and percent abundance of copper's other isotope?

Writing in Chemistry

88. The Standard Model of particle physics describes all of the known building blocks of matter. Research the particles included in the Standard Model. Write a short report describing the known particles and those thought to exist but not detected experimentally.

89. Individual atoms can be seen using a sophisticated device known as a scanning tunneling microscope. Write a short report on how the scanning tunneling microscope works and create a gallery of scanning tunneling microscope images from sources such as books, magazines, and the Internet.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

90. How is a qualitative observation different from a quantitative observation? Give an example of each. (Chapter 1)

91. A 1.0-cm^3 block of gold can be flattened to a thin sheet that averages 3.0×10^{-8} cm thick. What is the area (in cm^2) of the flattened gold sheet? A letter size piece of paper has an area of 603 cm^2. How many sheets of paper would the gold cover? (Chapter 2)

92. Classify the following mixtures as heterogeneous or homogeneous. (Chapter 3)

a. salt water
b. vegetable soup
c. 14-K gold
d. concrete

93. Are the following changes physical or chemical? (Chapter 3)

a. water boils
b. a match burns
c. sugar dissolves in water
d. sodium reacts with water
e. ice cream melts

STANDARDIZED TEST PRACTICE
CHAPTER 4

Use these questions and the test-taking tip to prepare for your standardized test.

1. An atom of plutonium _____ .

a. can be divided into smaller particles that retain all the properties of plutonium
b. cannot be divided into smaller particles that retain all the properties of plutonium
c. does not possess all the properties of a larger quantity of plutonium
d. cannot be seen using current technology

2. Neptunium's only naturally occurring isotope, $^{237}_{93}Np$, decays by emitting one alpha particle, one beta particle, and one gamma ray. What is the new atom formed from this decay?

a. $^{233}_{92}U$ **c.** $^{233}_{90}Th$
b. $^{241}_{93}Np$ **d.** $^{241}_{92}U$

3. An atom has no net electrical charge because _____ .

a. its subatomic particles carry no electrical charges
b. the positively charged protons cancel out the negatively charged neutrons
c. the positively charged neutrons cancel out the negatively charged electrons
d. the positively charged protons cancel out the negatively charged electrons

4. $^{126}_{52}Te$ has _____ .

a. 126 neutrons, 52 protons, and 52 electrons
b. 74 neutrons, 52 protons, and 52 electrons
c. 52 neutrons, 74 protons, and 74 electrons
d. 52 neutrons, 126 protons, and 126 electrons

5. Assume the following three isotopes of element *Q* exist: ^{248}Q, ^{252}Q, and ^{259}Q. If the atomic mass of *Q* is 258.63, which of its isotopes is the most abundant?

a. ^{248}Q **c.** ^{259}Q
b. ^{252}Q **d.** they are all equally abundant

Interpreting Tables Use the table to answer questions 6–8.

Characteristics of Naturally Occurring Neon Isotopes

Isotope	Atomic number	Mass (amu)	Percent abundance
^{20}Ne	10	19.992	90.48
^{21}Ne	10	20.994	0.27
^{22}Ne	10	21.991	9.25

6. Based on the table, an atom of neon found in nature would most likely have a mass of _____ .

a. 19.992 amu
b. 20.179 amu
c. 20.994 amu
d. 21.991 amu

7. In which of the neon isotopes is the number of neutrons the same as the number of protons?

a. ^{20}Ne
b. ^{21}Ne
c. ^{22}Ne
d. none of the above

8. The atomic mass of Ne is equal to _____ .

a. $\frac{19.922 \text{ amu} + 20.994 \text{ amu} + 21.991 \text{ amu}}{3}$
b. $\frac{1}{3}[(19.992 \text{ amu})(90.48\%) + (20.994 \text{ amu})(0.27\%) + (21.991 \text{ amu})(9.25\%)]$
c. $(19.992 \text{ amu})(90.48\%) + (20.994 \text{ amu})(0.27\%) + (21.991 \text{ amu})(9.25\%)$
d. 19.992 amu + 20.994 amu + 21.991 amu

9. Element X has an unstable nucleus due to an overabundance of neutrons. All of the following are likely to occur EXCEPT _____ .

a. element X will undergo radioactive decay
b. element X will eventually become a stable, nonradioactive element
c. element X will gain more protons to balance the neutrons it possesses
d. element X will spontaneously lose energy

10. The volume of an atom is made up mostly of _____ .

a. protons
b. neutrons
c. electrons
d. empty space

THE PRINCETON REVIEW **TEST-TAKING TIP**

Skip Around If You Can The questions on some tests start easy and get progressively harder, while other tests mix easy and hard questions. You may want to skip over difficult questions and come back to them later, after you've answered all the easier questions. This will guarantee more points toward your final score. In fact, other questions may help you answer the ones you skipped. Just be sure you fill in the correct ovals on your answer sheet.